

Acids and Alkalis

Teacher Notes



Acids and Alkalis is funded as part of the Reach and Teach educational programme supported by the Wolfson Foundation



THE
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Acids and Alkalis

Why focus on G&T and higher achievers?

Within the education system every child has the right to develop their learning so as to maximise their potential.

These exercises are designed to give students enthuse and enrich activities that although related to the curriculum are in fact taking the learning experience to the next level whilst also showing chemistry in a familiar context. This has been found to be a successful model for not only improving learning but also for raising levels of motivation. Higher achieving students can find the restraints of the standard curriculum to be demotivating leading to underachievement.

The different activities are designed to improve a number of skills including practical work/dexterity, thinking/analysis skills, literacy, research activities, use of models and teamwork. Students should also gain confidence through the activities and improve the ability to express themselves.

Some of the activities would appear to be complex for KS3 (year 9), however at this stage in their learning high achieving students are open to new concepts and are ready to explore issues without pre-conceptions. They are keen to link ideas and develop concepts and understanding. It can prove to be an uplifting experience.

Introduction

Understanding the background to acids and alkalis, and particularly focussing on the identification and quantification of pH, gives students a solid foundation for the investigation of chemical properties.

This programme is designed to develop students understanding of acids and alkalis as well as developing thinking and research skills.

Topic	Type of activity	Summary	Timing (mins)	KS3	KS4	KS5	Page
Making a pH indicator	Practical	Investigating what makes an indicator from natural substances. Using the indicators to test familiar substances.	35	√	√		7
Neutralisation circles	Practical	An introduction to the colours of universal indicator and the process of neutralisation.	10	√	√		9
An effervescent universal indicator reaction	Practical	A fun introduction of neutralisation using carbonates and the chemistry involved.	10	√	√		10

Investigation of indicators	Practical	A problem solving exercise investigating the varied range of indicators and their combination to form universal indicator.	20	√	√	√	11
Explaining acid strength	Misconception exercise	An exercise focussing upon learning and understanding that works through potential misconceptions.	Student centred		√	√	15
Acid-Base neutralisation	Practical	A quantitative approach to neutralisation, an acid-base titration on a microscale.	50	√	√	√	23
Acid revision map	Revision exercise	A revision of the topic leading to a concept map of acids	Student centred	√	√	√	28

The first activity looks at the use of naturally occurring dyes as a means of indicating if a substance is an acid or alkali, relating the tests to well known domestic material. This helps in learning as it puts the concepts within a familiar context. This is a format that is proven to accelerate learning.

This then leads into an activity that investigates the colour changes in universal indicator in a simple form and leads easily into neutralisation. Neutralisation can be visualised by the colour change. The third exercise builds on this, still looking at the colours of universal indicator, provides for extension in that a gas is produced. This can then be explored developing the reactions of acids with carbonates, testing of gases, and word/symbol equations. These are particularly good activities for key stage 3 and 4, although key stage 5 still delight at the effervescent rainbow!

Investigating indicators is good problem solving exercise for all key stages as it can be approached on several levels. At key stage 3 and 4 it is about the fact that indicators change colour in specific pH ranges, not necessarily pH 7. With this in mind some indicators can show acids as alkalis and vice versa. It also illustrates the foundation of universal indicator. At key stage 5 the idea of indicators as weak acids and the shifting of equilibria resulting in colour change can be explored.

The final two activities are intended to explore misconceptions and also to link together concepts and map these out. They act to extend and consolidate learning through independent and group work.

Aims and objectives

- The aims and objectives of these activities are:
- Developing questioning skills through problem solving.
- Exploring the use of models to expand understanding
- Develop practical skills and dexterity.

- Promote independent learning and research skills.

Chemistry topics:

- Acids and alkalis.
- Properties of indicators.
- The pH scale.
- Measuring pH.
- Quantitative analysis of pH.

These exercises can be used with key stages 3, 4, 5 as indicated on the *Possible Routes*.

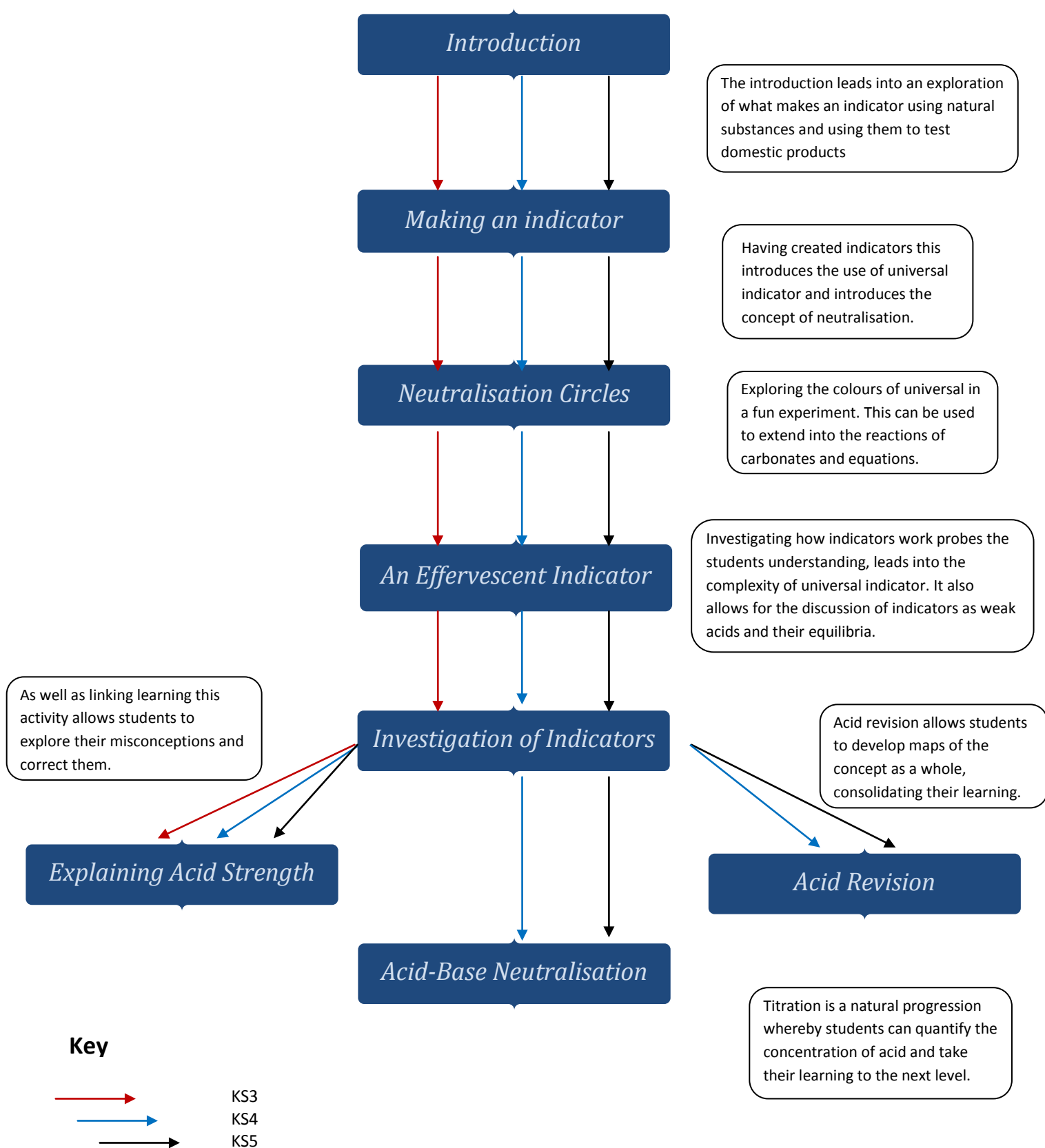
These activities have proved very successful with students at all levels who have followed the prescribed pathway and have been stimulated into further independent learning. Year 7 have even been comfortable with exploring some of the aspects.

At key stage 4 it has enhanced understanding of the whole basis of acids and their measurement.

These exercises further provide a reinforcement and revision tool for a number of topics from the A level syllabus.

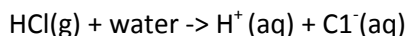
At all levels there is promotion of questioning skills, independent learning and research skills.

Possible routes

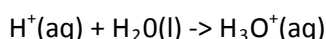


Detecting Acids and Alkalis

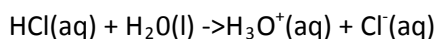
Acids are substances that dissociate in water to produce hydrogen ions. This idea was first put forward by the Swedish chemist Arrhenius in the nineteenth century. For example, although hydrogen chloride gas has a covalent structure, it dissolves in water to form:



As knowledge about atomic structure developed, chemists realised that the hydrogen ion was simply a proton, and it was highly unlikely to be stable on its own. It is now widely accepted that, from electrolysis reactions, to suggest that hydrogen ions combine with polar water molecules to form stable oxonium ions:



When hydrogen chloride dissolves in water, there is thought to be a chemical reaction:

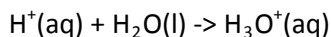


The same holds true for the other common acids: they all dissociate, to a greater or lesser extent, in water to form oxonium ions. This idea helps to explain why *acids only show their acidic nature when dissolved in water*. Traditionally, acids have been defined as substances that:

- produce hydrogen gas on reaction with metals;
- neutralise bases to give salt and water as the only products;
- release carbon dioxide when reacted with carbonates.

This information conveniently summarises the properties of most acids, but does not tell us anything about their structures and differences in reactivity. The evidence for the ionisation of acids (such as conductivity experiments) gives us an important extra property to include in a full definition – *acids provide hydrogen ions, $\text{H}^+(\text{aq})$, in solution. **Acids are proton donors***, and since *alkalis* are the chemical opposites of acids, they must be *proton acceptors*.

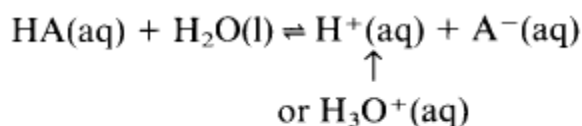
Here we should remind you about a common notation you will often see. It has just been said that the H^+ ion is thought to combine with water to form the more stable hydroxonium ion,



so the symbols H^+ , $\text{H}^+(\text{aq})$ and H_3O^+ can all refer to hydrogen ions in solution, and one or the other or even all may be used by different text books. For simplicity, this text will continue to use $\text{H}^+(\text{aq})$.

The Formation of Hydrogen Ions and the Acid Dissociation Constant

Experimentation shows us that some acids dissociate to a greater extent than others, *i.e.* more of their molecules produce H^+ ions and the equilibrium position is further to the right. For a general acid of formula HA:



By Le Chatelier's Principle, adding more water, *i.e.* diluting the acid, will *increase the ionisation* of the acid.

The acid dissociation constant should be:

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}][\text{H}_2\text{O}]}$$

but the concentration of the water present will hardly vary, so we can incorporate that constant into K_c , to give –

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} \text{ mol dm}^{-3}$$

where K_a is the acid dissociation constant.

The greater the degree of ionisation, the stronger the acid - the higher the value of K_a .

For acids of the same concentration, the greater the degree of ionisation of the acid, the more ions are present, and so the higher the conductivity.

Some typical values of acid dissociation constants are given in table below:

| *Acid dissociation constants for some common organic acids*

<i>Acid</i>	<i>Formula</i>	<i>Acid dissociation constant, K_a (mol dm⁻³)</i>
Ethanoic	CH ₃ COOH	1.8×10^{-5}
Trichloroethanoic	CCl ₃ COOH	2.3×10^{-1}
Benzoic	C ₆ H ₅ COOH	6.3×10^{-5}
Methanoic	HCOOH	1.6×10^{-4}

Why is the value of K_a for trichloroethanoic acid more than 10000 times larger than that for ethanoic acid?

Chlorine is an electronegative atom; it pulls electrons towards its nucleus. This effect is felt throughout the molecule, and thus the O-H bond is weakened with the result that hydrogen ions are more easily formed. More hydrogen ions in solution means a stronger acid.

The values of K_a for the common inorganic acids, such as sulfuric and nitric acids, are of course very high because the acids are fully ionised in dilute solution. They are called strong acids.

You will also come across references to $\text{p}K_a$,

$$\text{p}K_{\text{a}} = -\log K_{\text{a}}$$

Using this definition - **the smaller the value of $\text{p}K_{\text{a}}$, the stronger the acid.**

It is useful for students to realise that indicators are coloured dyestuffs, and one of the most effective ways of achieving this is to make extracts of naturally occurring indicators. It is often surprising for students to realise what dramatic colour changes can be achieved with such things as the extract from red pelargonium petals or elderberries.

This can be further placed in context by using the indicators on everyday items such as household cleaners and domestic products.

Activity 1: Making a pH indicator

A pH indicator is a substance which has one colour when added to an acidic solution and a different colour when added to an alkaline solution. In this experiment pupils make an indicator from red cabbage. It is also possible to use other materials such as beetroot, berries or flower petals.

Lesson organisation

The experiment is in two parts. The first part involves boiling some red cabbage in water. In the second part the students test their indicator. Between the two parts the mixture must be allowed to cool. The first part takes about 10 to 15 minutes. The cooling takes about 15 minutes and the testing less than 5 minutes.

The cooling period could be used as an opportunity to discuss the background to the experiment – see Teaching notes below.

Apparatus and chemicals

Eye protection for all

Each working group will require:

Beaker (250 cm³)

Bunsen burner

Tripod

Gauze

Heat resistant mat

Test-tubes, 3 (see note 1)

Test-tube rack

Dropper pipette

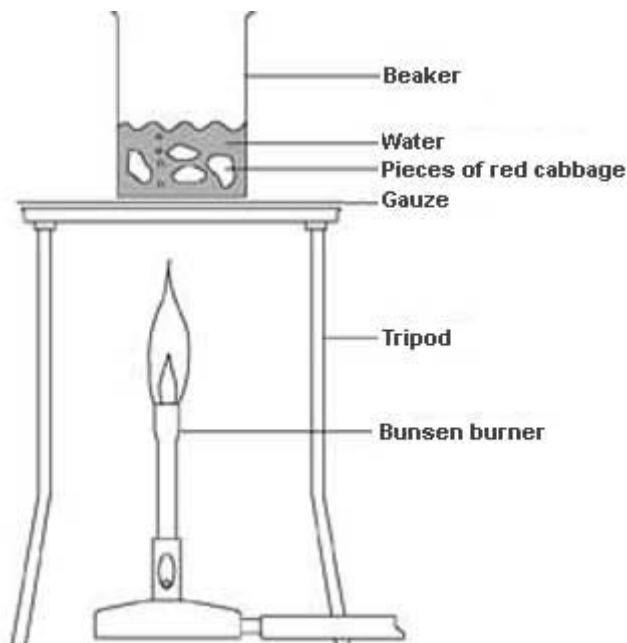
Several pieces of red cabbage

Access to (see notes 2 and 3):

Dilute hydrochloric acid, 0.01 mol dm^{-3} (**Low hazard** at this concentration)

Sodium hydroxide solution 0.01 mol dm^{-3} (**Low hazard** at this concentration)

De-ionised or distilled water



Technical notes

Dilute hydrochloric acid, 0.01 mol dm^{-3} (**Low hazard** at concentration used). Refer to CLEAPSS Hazard 47A

Sodium hydroxide solution 0.01 mol dm^{-3} (**Low hazard** at concentration used). Refer to CLEAPSS Hazard 91

1 Small test-tubes of capacity about 10 cm^3 are ideal.

2 Each group of students will need access to the hydrochloric acid and sodium hydroxide solutions. Dropper bottles are ideal. Alternatively small beakers (100 cm^3) with dropper pipettes could be used. Students need to be able to pour the acid and alkali solutions easily and safely into test-tubes.

3 Provide similar containers for de-ionised or distilled water. Label the containers 'Acid', 'Alkali' and 'Water'.

Procedure

HEALTH & SAFETY: Wear eye protection throughout. Consider clamping the beaker. Students should remain standing whilst water is boiling.

- a) Boil about 50 cm^3 of water in a beaker.
- b) Add 3 or 4 small (5 cm) pieces of red cabbage to the boiling water.

- c) Continue to boil the red cabbage in the water for about 5 minutes. The water should turn blue or green.
- d) Turn off the Bunsen burner and allow the beaker to cool for a few minutes.
- e) Place 3 test-tubes in a test-tube rack. Half-fill one of the test-tubes with acid, one with alkali, and one with distilled or de-ionised water. Label the test-tubes.
- f) Use a dropper pipette to add a few drops of the cabbage solution to each test-tube. Note the colour of the cabbage solution in each of the three test-tubes.

Teaching notes

Discussion points could include any or all of the following.

Many plant colouring materials in berries, leaves and petals act as indicators.

Some of these will not dissolve in water easily. A solvent other than water (e.g. ethanol) could be used, but it may be flammable. Discuss how the risk of fire can be reduced by using a beaker of hot water to heat the mixture.

Possible variations on this experiment might include using beetroot, blackberries, raspberries, copper beech leaves, or onion skins in place of the red cabbage.

A good introduction to the measurement of pH and the colours of universal indicator is a simple practical, neutralisation circles.

Activity 2: 'Neutralisation circles'

Teacher notes

Description

Drops of dilute acid and alkali are placed a few centimetres apart on a sheet of filter paper and allowed to spread out until they meet. A few drops of Universal indicator are then placed over the moist area of the filter paper and a band of colours showing the range of colours of the Universal indicator is seen on the paper.

This experiment will take around 10 minutes.

Procedure

Students are given a piece of filter paper and asked to draw on it in pencil two circles about 1 cm in diameter and about 2 – 3 cm apart which they label 'acid' and 'alkali' respectively. The filter paper is then placed on a white tile and students use dropping pipettes to place a few drops of the appropriate solution in each circle. The concentrations of the acid and alkali are not critical but they should be approximately the same. The solution will begin to spread out on the filter paper. The students wait for a few minutes until the solutions have soaked through the filter paper towards each other and have met. Students then place drops of Universal Indicator solution on the area of the filter paper where the acid and alkali have met

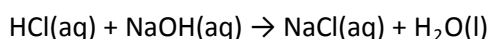
and reacted. A 'rainbow' is produced showing the range of colours produced by universal indicator.

Students can dry the filter papers and stick them into their notebooks.

Teaching notes

This experiment is quicker, simpler and safer than the traditional method of illustrating neutralisation by titrating acid with alkali using a burette. It also uses more familiar equipment (a dropping pipette rather than a burette), uses little of the reagents and has the advantage of producing a permanent record of the colour changes.

The reaction is



Another simple version is the effervescent rainbow, which again shows the colours of universal indicator but can be further developed to introduce balanced equations and discussions as to how the students could identify the products. For example the use of limewater to test for carbon dioxide.

Activity 3: An effervescent universal indicator 'rainbow'

Teacher notes

Description

Sodium carbonate solution is added to a burette containing a little hydrochloric acid and Universal Indicator. The two solutions react, with effervescence, and the liquid in the burette shows a 'rainbow' of all the colours of Universal Indicator from red through orange, yellow, green and blue to purple.

The practical takes about 5 – 10 minutes.

Method

Clamp the burette vertically. Add about 0.5 cm³ of the universal indicator solution followed by about 10 cm³ of the hydrochloric acid (**irritant**) to give a clearly visible red colour. Now add about 20 cm³ of the sodium carbonate solution. Insert a loose plug of cotton wool in the top of the burette. The sodium carbonate and hydrochloric acid react, with effervescence, and the burette will be filled with liquid showing a 'rainbow' of all the colours of universal indicator from red through orange, yellow, green and blue to purple.

Visual tips

A white background will show the colours to best advantage.

Health & Safety

Wear eye protection.

An effervescent universal indicator 'rainbow'

Technician notes

Apparatus

Each group will need:

- Eye protection
- A 50 cm³ burette
- A retort stand with boss and clamp
- Cotton wool plug

Chemicals

Each group will need:

- A few cubic centimetres of Universal Indicator solution
- About 10 cm³ hydrochloric acid solution (2 mol dm⁻³) (**irritant**)
- About 20 cm³ sodium carbonate solution (1 mol dm⁻³) (solid is an **irritant**)

Few indicators change colour at pH=7, this class practical develops students thinking skills and problem solving skills as it focuses on the limitations of different indicators. In this indicator practical students measure the pH of a range of solutions and attempt to explain what appear to be erroneous results – for example sodium hydrogen carbonate showing as an acid. At the end of the experiment they make their own basic universal indicator solution that they use to retest all the samples.

Activity 4: Investigation of indicators

Teacher notes

Indicators are substances that, when added to solutions of different pH, change colour. You are provided with three: bromothymol blue, methyl orange and phenolphthalein (highly flammable).

1. Wearing goggles, begin your investigation by finding what colour each of these indicators turns when added to an acid or an alkali. To do this pour about 2 cm depth of dilute hydrochloric acid into a test tube and add two drops of one of the indicators to find out what colour it turns with acids. Then pour about 2 cm depth of sodium hydroxide solution into another test tube and add two drops of the indicator to find out what colour it turns in alkalis. Repeat this using clean test tubes for the other two indicators. Write the colours you observe in the spaces below:

INDICATOR	COLOUR IN ACIDS	COLOUR IN ALKALIS
Bromothymol Blue	yellow	blue
Methyl Orange	red	yellow
Phenolphthalein	colourless	pink

2. You are now going to test two solutions with the three indicators and try to draw conclusions about them from the colours shown by the indicators. The solutions with which you are provided are **boric acid solution** and **sodium bicarbonate solution**. To investigate these, pour about 2 cm depth of one of the solutions into each of three clean test tubes and add two drops of an indicator to one test tube. Add two drops of another indicator to the second tube and two drops of the third indicator to the third tube. From the colours shown by the indicator in each case decide whether it is indicating ACID or ALKALI. Repeat this part of the investigation for the other solution. Now record your observations and conclusions in the following table:

	Results with Boric acid solution	Results with Sodium bicarbonate solution
Colour with Bromothymol Blue	yellow	blue
Acid or alkali shown by Bromothymol Blue?	acid	alkali
Colour with Methyl Orange	yellow	yellow
Acid or alkali shown by Methyl Orange?	alkali	alkali
Colour with Phenolphthalein	colourless	colourless
Acid or alkali shown by Phenolphthalein?	acid	acid

Is there anything odd you notice in your results in the above table? YES/NO (Delete the answer which does not apply) *YES!*

Explain your answer:

Boric acid appears to be alkali according to methyl orange

Sodium bicarbonate appears to be acid according to phenolphthalein

3. To demonstrate the action of these indicators further, you are going to test some solutions with a **mixture** of the indicators. First prepare this mixture by mixing in a test tube:

- 10 drops of Bromothymol blue,
- 5 drops of methyl orange and
- 5 drops of phenolphthalein.

Mix them well and, using a teat dropper, add two drops of this mixed indicators solution to test tubes containing 2 cm. depth of each of the following substances:

- Dilute hydrochloric acid,
- Boric acid solution,
- tap water,
- sodium bicarbonate solution and
- sodium hydroxide solution.

Record your observations in the table below:

SUBSTANCE	COLOUR SHOWN BY MIXED INDICATORS SOLUTION
Dilute hydrochloric acid	RED
Boric acid solution	YELLOW
Tap water	GREEN
Sodium bicarbonate solution	BLUE
Sodium hydroxide solution	PURPLE

Have you ever seen a series of colours similar to this shown by an indicator solution? YES/NO
(Delete the answer which does not apply) *yes*

Comment on the colours you see in this part of the investigation and try to explain your results:

Colours of the rainbow/colours shown by universal indicator/ different indicators change at different pHs/Boric acid only weak acid – not strong enough to turn methyl orange to red/Sodium bicarbonate only weak alkali – not strong enough to turn phenolphthalein pink

	Bromothymol Blue	Methyl Orange	Phenolphthalein	Mixed
HCl	<i>yellow</i>	<i>red</i>	<i>colourless</i>	<i>Red/orange</i>
Boric acid	<i>yellow</i>	<i>yellow</i>	<i>colourless</i>	<i>Yellow</i>
Tap water	<i>green</i>	<i>yellow</i>	<i>colourless</i>	<i>Yellow/Green</i>
NaHCO ₃	<i>blue</i>	<i>yellow</i>	<i>colourless</i>	<i>Blue/green</i>
NaOH	<i>blue</i>	<i>yellow</i>	<i>pink</i>	<i>Purple</i>

Technician notes

Each group will require

2 spotting tiles

1 test tube + a rack

Dropper bottles of

- **Methyl orange** indicator solution
- **Bromothymol blue** indicator solution
- **Phenolphthalein** indicator solution

Bottles of

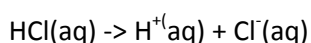
- dilute hydrochloric acid (~0.1M) (**low hazard**)
- Sodium hydroxide solution (~0.1M) (**irritant**)
- Boric acid solution (1 teaspoonful – dissolve in 250ml of hot distilled water, allow to cool and filter if any crystallizes out) (solid- **toxic**, solution – **low hazard**) See CLEAPSS *Hazcard* 14.
- Sodium hydrogencarbonate solution (1 teaspoonful per 250 ml, then saturate with carbon dioxide by passing gas through from a cylinder for a few minutes. If you do not have a cylinder of carbon dioxide, make up the solution COLD using fizzy mineral water! Do not heat this.)

5 disposable plastic droppers.

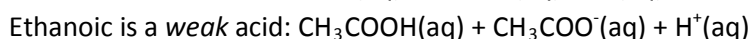
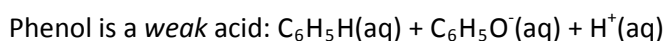
Access to tap water

STRONG AND WEAK ACIDS

Hydrochloric acid is said to be a strong acid as it is completely dissociated in dilute solution:



Ethanoic acid and phenol are weak acids as they are never completely ionised, however dilute the solution.



Many weak acids occur naturally and are organic in origin. For instance:

- methanoic acid (which used to be called formic acid) is present in the sting of ants
- ethanoic acid (formerly called acetic acid) is in vinegar - it has a characteristic and very pungent smell
- butanoic acid gives rancid butter and cheese its characteristic odour

N.B. Do not confuse the terms **STRONG** and **CONCENTRATED**.

Strong acids are completely dissociated into ions in dilute solutions;

concentrated acids (or any concentrated solution) contain several moles of substance per dm^3 of solution. Ordinary 'lab' concentrated hydrochloric acid contains approximately 10 mol dm^{-3} of $\text{HCl}(\text{aq})$ and concentrated sulfuric acid contains about 18 mol dm^{-3} $\text{H}_2\text{SO}_4(\text{aq})$.

The concept of 'strong' and 'weak' may be applied to any electrolyte to indicate the degree to which they dissociate (break up) into ions.

THE USE OF THE pH SCALE

The acidity, or alkalinity of a solution, is frequently assessed using universal indicator paper or solution, which is a mixture of a variety of indicators that go through a spectrum of colours as the acidity of a solution changes. In the table below, pH values have been matched to H^+ ion concentration and the usual colour of universal indicator.

2 Relation between pH, $[\text{H}^+]$ and the colour of universal indicator

<i>pH scale</i>	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
$[\text{H}^+]$	1	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}		10^{-7}			10^{-10}	10^{-11}			10^{-14}
<i>Colour of universal indicator</i>	Red		Orange		Yellow		Green			Blue		Purple			
	Acid						Neutral				Alkaline				

The following activity helps students consolidate their learning by focussing on the explanation of acid strength. This is particularly useful for key stages 4 and 5.

Activity 5: Explaining acid strength

One definition of an acid is that it dissolves in water to give hydrogen ions (H^+).

In fact the hydrogen ion (H^+) will associate with a water molecule to form H_3O^+ .

One way to write the equation for an acid 'HA' dissolving in water is:



The A in HA does not stand for a particular element, but for the 'acid radical' part of the molecule. So, for example, in hydrochloric acid 'HA' would be HCl, and 'A⁻' would be Cl^- , whilst in ethanoic acid 'HA' would be CH_3COOH , and 'A⁻' would be CH_3COO^- .

Acids (and alkalis) can be described as 'strong' or 'weak', and as 'concentrated' or 'dilute'.

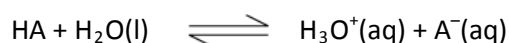
1. What is the difference between a strong acid and a weak acid?

2. What is the difference between a concentrated acid and a dilute acid?

3. If you could see the particles (molecules, ions etc) in an acidic solution, how would you decide whether it was a solution of a strong acid or a solution of a weak acid?

Classifying acid solutions

One way to write the equation for an acid 'HA' dissolving in water is:

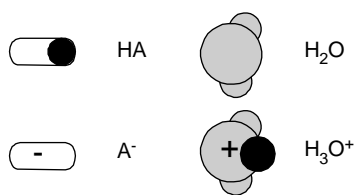


On the following pages are some diagrams of acidic solutions.

Scientific diagrams are always simplifications designed to highlight some aspects of the system represented. The diagrams in this exercise show simplifications of real solutions.

For example, the concentration of acids varies over many orders of magnitude, and an accurate diagram of a very dilute solution would need to show many thousands of water molecules for each $\text{H}^+(\text{aq})$ ion.

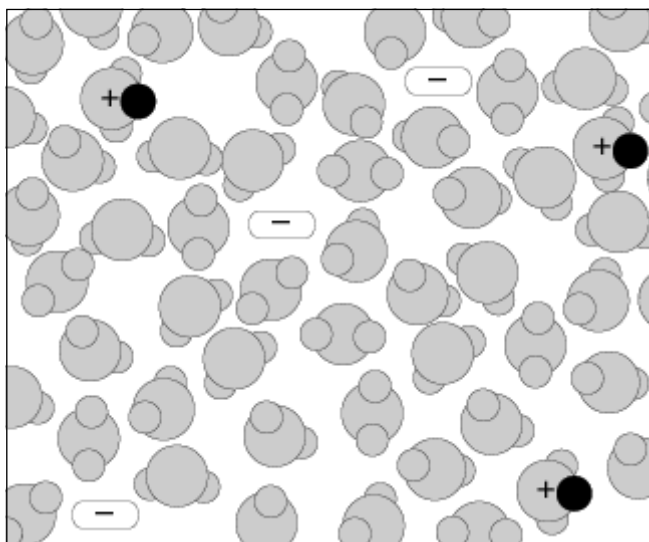
Only four types of particle are shown in these diagrams. The following key is used to distinguish between the different particles:



The size (and shape) of acid molecules varies greatly, and they are often much larger than a water molecule.

Look carefully at the four diagrams on the following pages, and see if you can tell what the differences between them are meant to indicate.

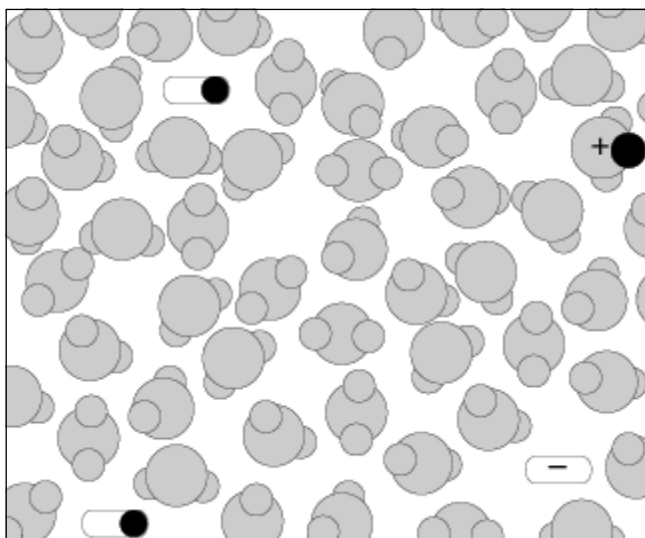
Diagram 1



1. What types of particles are shown in the solution represented in this diagram?

2. How would you describe this solution?

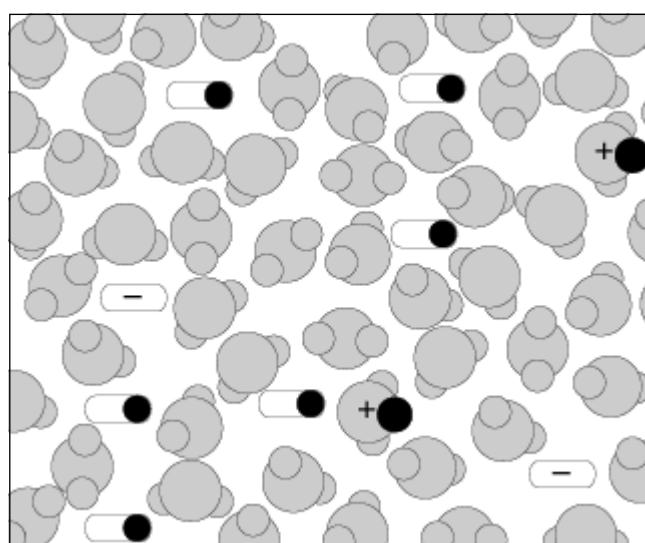
Diagram 2



3. What types of particles are shown in the solution represented in this diagram?

4. How would you describe this solution (compared to diagram 1)?

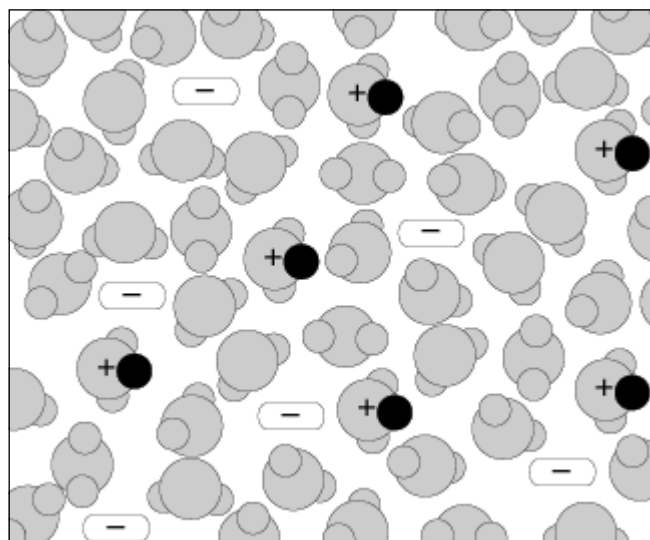
Diagram 3



5. What types of particles are shown in the solution represented in this diagram?

6. How would you describe this solution (compared to diagrams 1 and 2)?

Diagram 4

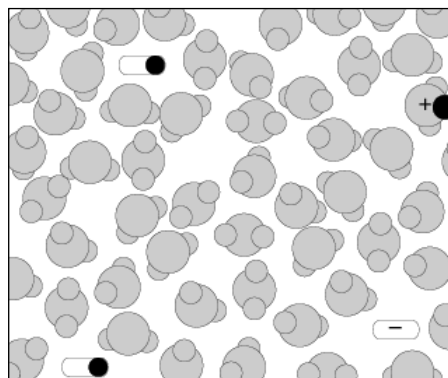
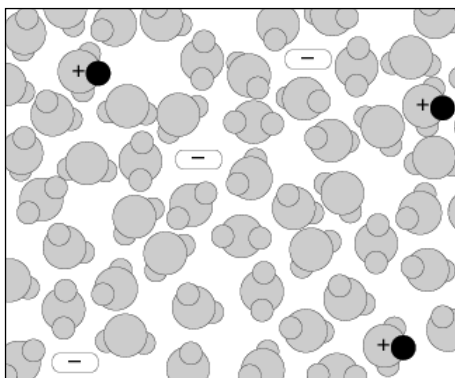


7. What types of particles are shown in the solution represented in this diagram?

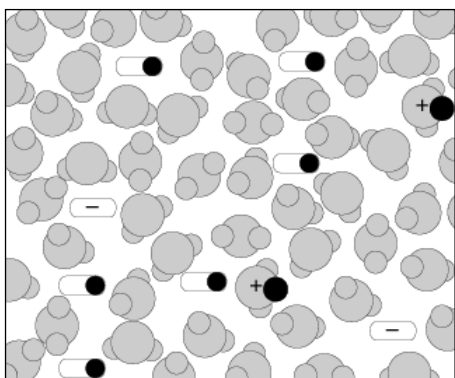
8. How would you describe this solution (compared to diagrams 1-3)?

9. The four diagrams you were asked to consider are reproduced in miniature below.

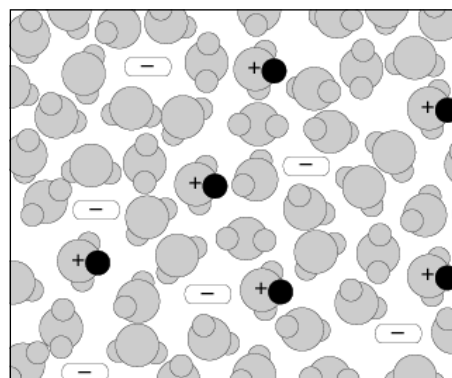
2.



3.



4.



The diagrams are meant to represent a concentrated solution of a strong acid, a dilute solution of a strong acid, a concentrated solution of a weak acid and a dilute solution of a weak acid. Use the table below to show which diagram is meant to represent each of the four solutions – write the number of the appropriate diagram in each box.

	Strong	Weak
Concentrated		
Dilute		

The Measurement of pH, and the pH Scale

CALCULATING VALUES OF pH

The strength of an acid is indicated by its pH: this is a sort of 'upside down' measure of the concentration of hydrogen ions in solution. As the pH goes up, so the concentration of hydrogen ions in solution, $[H^+]$, goes down. The 'p' part of pH comes from the German word 'potenz' meaning power.

In pure water and in neutral aqueous solutions such as sodium chloride, any hydrogen or hydroxide ions present come *only* from the ionisation of water. Hence:

$$[H^+] = [OH^-] = 1 \times 10^{-7} \text{ mol dm}^{-3}$$

for convenience this is often written as 10^{-7} .

Let us stop for a moment and consider the size of the numbers we are using. Every time that we want to state the strength of an acid we need, at best, to express the concentration of H^+ ions in standard index form or write out a string of noughts. If, however, we use the base 10 logarithms (logs) of very large or very small numbers they become much more manageable.

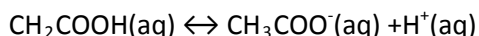
By definition $\log_{10} 0.1 = \log_{10} 10^{-1} = -1$
and $\log_{10} 0.001 = \log_{10} 10^{-3} = -3$, etc.
And don't forget $\log_{10} 1 = \log_{10} 10^0 = 0$

- pH is defined as the negative logarithm to base 10 (written as log) of the hydrogen ion concentration, $[H^+]$, in a solution. This convention means that chemists can use small positive numbers to express $[H^+]$ instead of working in powers of ten.
 $pH = -\log [H^+]$
so the pH of pure water = $-\log [10^{-7}]$ which is 7
so *any* solution with pH 7 is neutral.
- On your calculator input **0.000 0001** then press **[log] [=]**
What is the result?
- You can also use the exponential EXP button on the calculator:

[1] [EXP] [+/-] [7] [log] [=]

e.g.

- In 1 dm^3 of 0.1 mol dm^{-3} nitric acid (or 0.1 M) nitric acid of concentration 0.1 mol dm^{-3} therefore has a pH of 1.
- Dilute nitric acid is *completely ionised*, there are 0.1 or 10^{-1} moles of hydrogen ions - - it has a low pH.
- $\text{HNO}_3(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{NO}_3^-(\text{aq})$
- In 1 dm^3 of 0.1 mol dm^{-3} ethanoic acid, however, there are only 0.001 or 10^{-3} moles of hydrogen ions.
- Ethanoic acid of concentration 0.1 mol dm^{-3} has a pH of 3.
- This is because ethanoic acid is only *partially ionised*:



CALCULATING THE pH OF ALKALINE SOLUTIONS

It is easy to see how the pH of an acid is worked out, because by definition an acid is a proton (H^+) donor; but what about alkalis?

A solution of sodium hydroxide contains 0.1 moles of NaOH per dm^3 .

- It has a *concentration* of 0.1 mol dm^{-3} .
- Sodium hydroxide is a strong alkali - it is completely ionised.
- So $[\text{OH}^-]$ ions = 10^{-1} mol dm^{-3} (remember that $[\text{X}]$ means concentration of X, in mol dm^{-3}).
- But, for an aqueous solution, the ionic product of water is constant.
- $[\text{H}^+] \times [\text{OH}^-] = 10^{-14}$ mol dm^{-3}

$$\text{so } [\text{H}^+] = \frac{10^{-14}}{[\text{OH}^-]} = \frac{10^{-14}}{10^{-1}} = 10^{-13} \text{ mol dm}^{-3}$$

- $\text{pH} = -\log[\text{H}^+] = -\log 10^{-13} = 13$

Therefore: Sodium hydroxide solution of concentration 0.1 mol dm^{-3} has a pH of 13.

Generally:

Neutral solutions $[\text{H}^+] = [\text{OH}^-] = 10^{-7}$ mol dm^{-3} and $\text{pH} = 7$ at 25°C (298 K)

Acidic solutions $[\text{H}^+] > [\text{OH}^-]$, and $\text{pH} < 7$ at 25°C

Alkaline solutions $[\text{H}^+] < [\text{OH}^-]$, and $\text{pH} > 7$ at 25°C

Try these calculations

1. What is the pH of a solution containing $0.0001(10^{-4})$ moles H^+ ions?
2. What is the pH of a solution of hydrochloric acid $[\text{HCl}(\text{aq})]$ containing 0.02 moles H^+ ions per dm^3 ?
3. What is the pH of a solution containing 5×10^{-4} mol dm^{-3} H^+ ions?
4. What will be the pH of 1.7×10^{-5} mol dm^{-3} $\text{HCl}(\text{aq})$?
5. A solution of a strong base contains 0.01 mol dm^{-3} hydroxide ions (OH^-). What is the pH of the solution?

Answers

6. $\text{pH} = 4$
7. $\text{pH} = 1.698 = 1.7$
8. $\text{pH} = 3.3$
9. $\text{pH} = 4.769 = 4.8$
10. $\text{pH} = 12$

Neutralisation

The concept of neutralisation has been touched upon earlier in the text. It is a good extension to review this again from a more quantitative approach. Undertaking this microscale neutralisation practical gives the opportunity to link the measurement of pH with balanced equations, moles, analysis and evaluation. This microscale version has the benefit of being both quick and effective, as well as limiting the amount of equipment and materials used.

Activity 6: Acid-base neutralisation – a microscale titration

A microscale titration apparatus is prepared from pipettes, a syringe and some rubber or plastic tubing. This is then used to carry out a titration by filling the 'burette' with hydrochloric acid and placing 1 cm³ of sodium hydroxide solution in a 10 cm³ beaker. The aim is to calculate the exact concentration of the sodium hydroxide solution.

Lesson organisation

Microscale techniques are a fairly recent innovation in school chemistry, but most students take readily to them. Manipulative skills are important, and students need to be capable of careful manipulation to carry this out successfully. Students also need to be familiar with the mole concept, and capable of performing the calculations from the results of the experiment.

On such a small scale, safety issues are minimal. Similarly, the time taken to carry out a titration should be much reduced as the volumes being reacted are so small. It should be possible for a class to carry out the practical work and calculations in a one-hour session.

Apparatus and chemicals

Eye protection

Each working group will require:

Microscale titration apparatus (see note 1):

Graduated glass pipette (2 cm³)

Pipette (1 cm³) + pipette filler to fit (or a 1 cm³ plastic syringe)

Plastic syringe (10 cm³)

Fine-tip poly(ethene) dropping pipette (see note 2)

Small lengths of rubber, plastic or silicone tubing

Beakers (10 cm³), 2

Clamp stand with two bosses and clamps

Hydrochloric acid, 0.10 mol dm⁻³ (**Low hazard** at this concentration), about 10 cm³

Sodium hydroxide, approx 0.1 mol dm⁻³ (**Irritant** at this concentration), about 10 cm³ (note 3)

Phenolphthalein indicator solution (**Highly flammable**), a few drops

Technical notes

Hydrochloric acid (**Low hazard** at concentration used) Refer to CLEAPSS Hazcard 47A and CLEAPSS Recipe Book sheet 43

Sodium hydroxide (**Irritant** at concentration used) Refer to CLEAPSS Hazcard 91 and CLEAPSS Recipe Book sheet 85

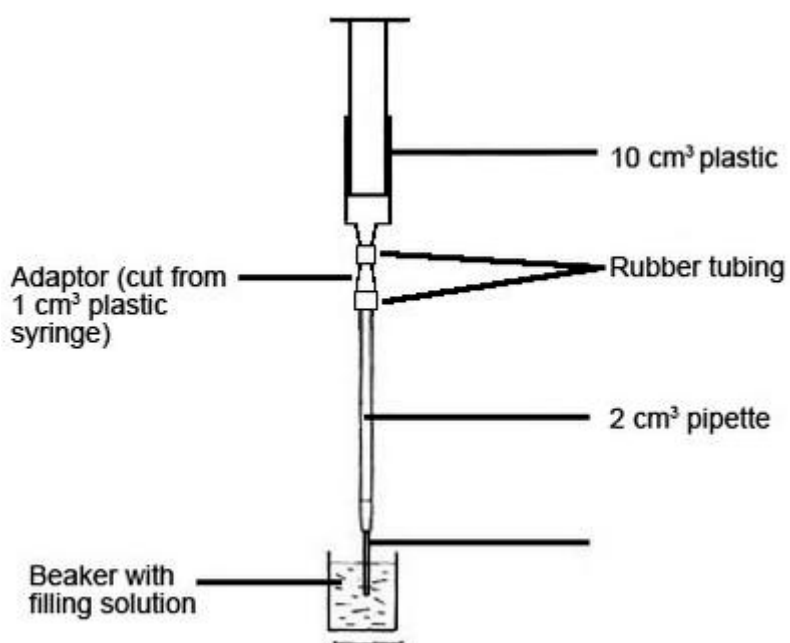
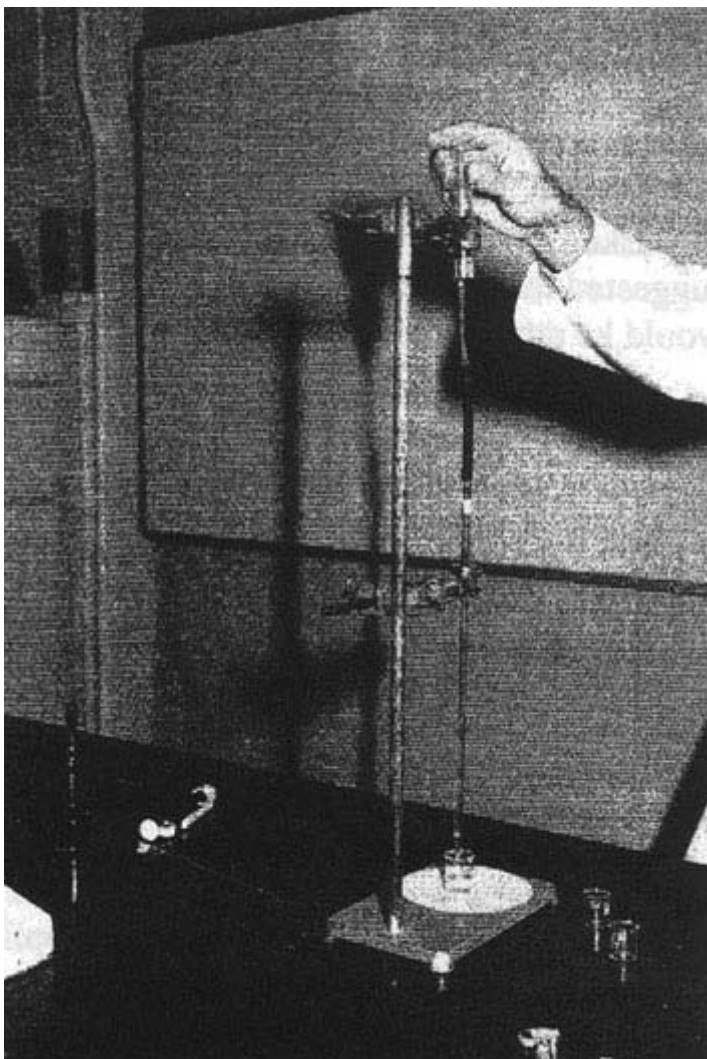
Phenolphthalein indicator solution (**Highly flammable**) Refer to CLEAPSS Hazcard 32 and CLEAPSS Recipe Book sheet 43

1 The microscale titration apparatus replaces the normal burette. To make the microscale titration apparatus, cut the tip end off a fine-tip poly(ethene) dropping pipette and push the tip carefully onto the end of a 2 cm³ graduated glass pipette. Clamp a plastic syringe, 10 cm³ capacity, above the adapted pipette, as shown in the picture, and connect the two with rubber, plastic, or silicone tubing. Because the diameters of the syringe nozzle and of the top of the pipette may be quite different, two pieces of tubing, one to fit each end, will probably be needed; these can then be joined by an adaptor. A suitable adaptor can be made by cutting the lower end off a 1 cm³ plastic syringe, such that the syringe body diameter fits the wider tubing, and the syringe tip fits the narrower tubing. See diagram and photograph.

It is possible for students to build their own microscale titration apparatus from supplied components, but this is likely to take the students more time than the titration itself! For that reason, it is probably preferable to prepare a class set of these in advance.

2 A suitable poly(ethene) dropping pipette would be fine-tip standard, non-sterile, 3.3 cm³ capacity, such as those available from Sigma-Aldrich.

3 Students are to calculate the concentration of the sodium hydroxide solution so the bottle should not be labelled with the exact concentration.

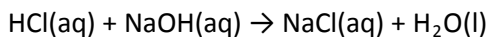


Procedure

HEALTH & SAFETY: Wear eye protection

- Clamp the microscale titration apparatus securely in position as in photograph and push the syringe plunger completely down.
- Fill the apparatus with 0.10 mol dm^{-3} hydrochloric acid as follows. Put about 5 cm^3 of the acid in a 10 cm^3 beaker and place the tip of the apparatus well down into the solution. Raise the syringe plunger slowly and gently, making sure no air bubbles are drawn in. Fill the pipette exactly to the zero mark. Release the plunger; the level should remain steady.
- Use the 1 cm^3 pipette and pipette filler to transfer exactly 1.0 cm^3 of the sodium hydroxide solution into a clean 10 cm^3 beaker.
- Add one drop (no more!) of phenolphthalein indicator solution to the sodium hydroxide solution.
- Adjust the position of the microscale titration apparatus so that the tip is just below the surface of the sodium hydroxide and indicator solution in the beaker
- Titrate the acid solution into the alkali by pressing down on the syringe plunger very gently, swirling to allow each tiny addition to mix and react before adding more.
- Continue until the colour of the indicator just turns from pink to permanently colourless.
- Record the volume of hydrochloric acid added at that point.
- Repeat the titration until you get reproducible measurements – that is, the volume required is the same in successive titrations.
- Calculate the concentration of the sodium hydroxide solution as follows.

The equation for the neutralisation reaction is:



From the equation you can see that one mole of hydrochloric acid reacts with one mole of sodium hydroxide.

1 What was the reliable value for the volume of hydrochloric acid solution needed? Let us call this value $V \text{ cm}^3$.

2 Calculate the number of moles of hydrochloric acid in this volume using the formula: $V/1000 \times C$, where C is the concentration of the hydrochloric acid in mol dm^{-3} .

3 How many moles of sodium hydroxide were therefore present in the original 1 cm^3 of sodium hydroxide solution placed in the beaker?

4 Now calculate how many moles of sodium hydroxide would have been present in 1000 cm^3 . This is the concentration of the sodium hydroxide solution in mol dm^{-3} .

Teaching notes

This microscale technique minimises apparatus and chemical requirements, and takes less time to perform than titration on the usual scale. Although the solutions used do present minor hazards, the use of such small quantities reduces risks from those hazards to very low levels. Students should nevertheless take all usual precautions in handling these solutions. The main risk is from misuse of the syringe or pipettes, especially if containing hazardous solutions.

The technique also makes the point that quantitative chemical experimentation does not always have to be performed on the traditional 'bucket' scale at school level.

Web links

Other examples of school use of microscale titration can be found at:

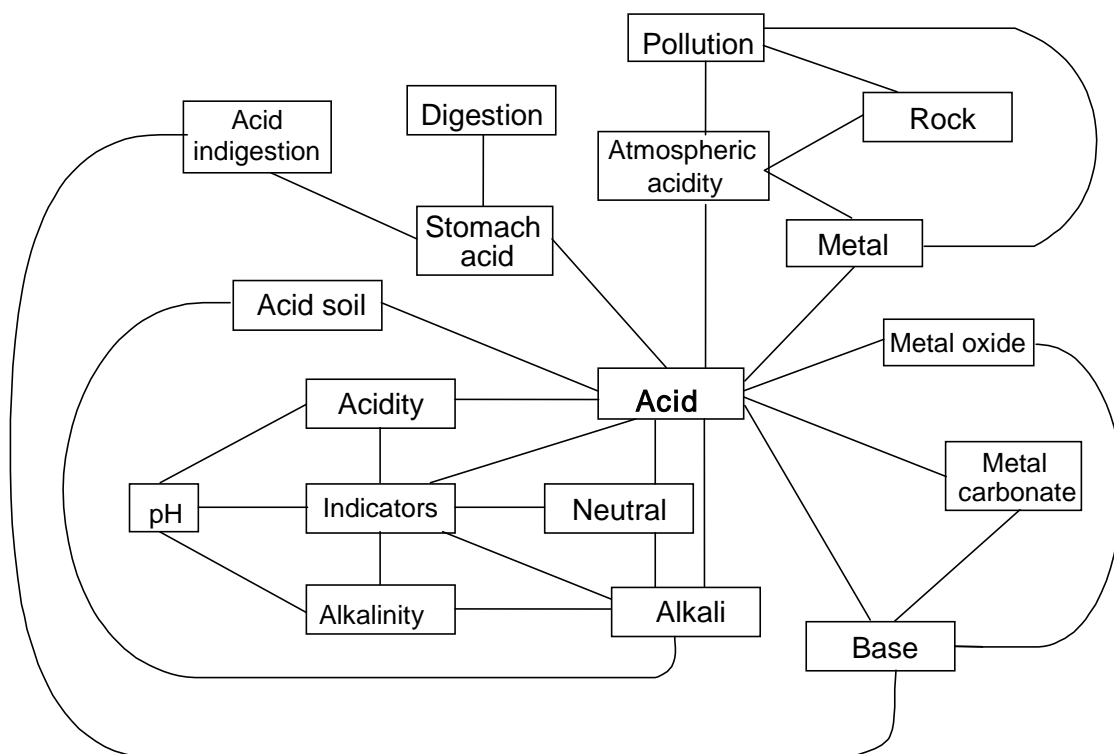
<http://dwb.unl.edu/Chemistry/MicroScale/MScale18.html>

The next link is to a commercial site which can supply ready-made microscale titration kits:

http://www.sargentwelch.com/product.asp_Q_pn_E_WL5261-55_EA_A_Somerset+Titration+Kit_E_

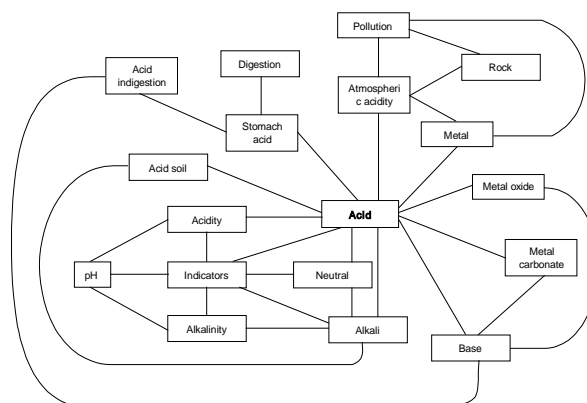
The Acid Revision Map pulls together all the students learning on acids. It is suitable for all key stages but the explanations that the students give may reflect their experience to date. It covers aspects that have not been directly part of this pack but this is a perfect opportunity for students to undertake individual research and learning in order to be able to answer the questions. Alternatively, it provides an opportunity for students to work together, pooling their knowledge and understanding, developing their group work and communication skills.

Activity 7: Acid revision map



Labelling the revision map

You have been given a copy of the acid revision map. This shows some of the important ideas you may have met when you studied acids and bases in your science class. Each line on the map stands for an idea that could be put into a sentence.



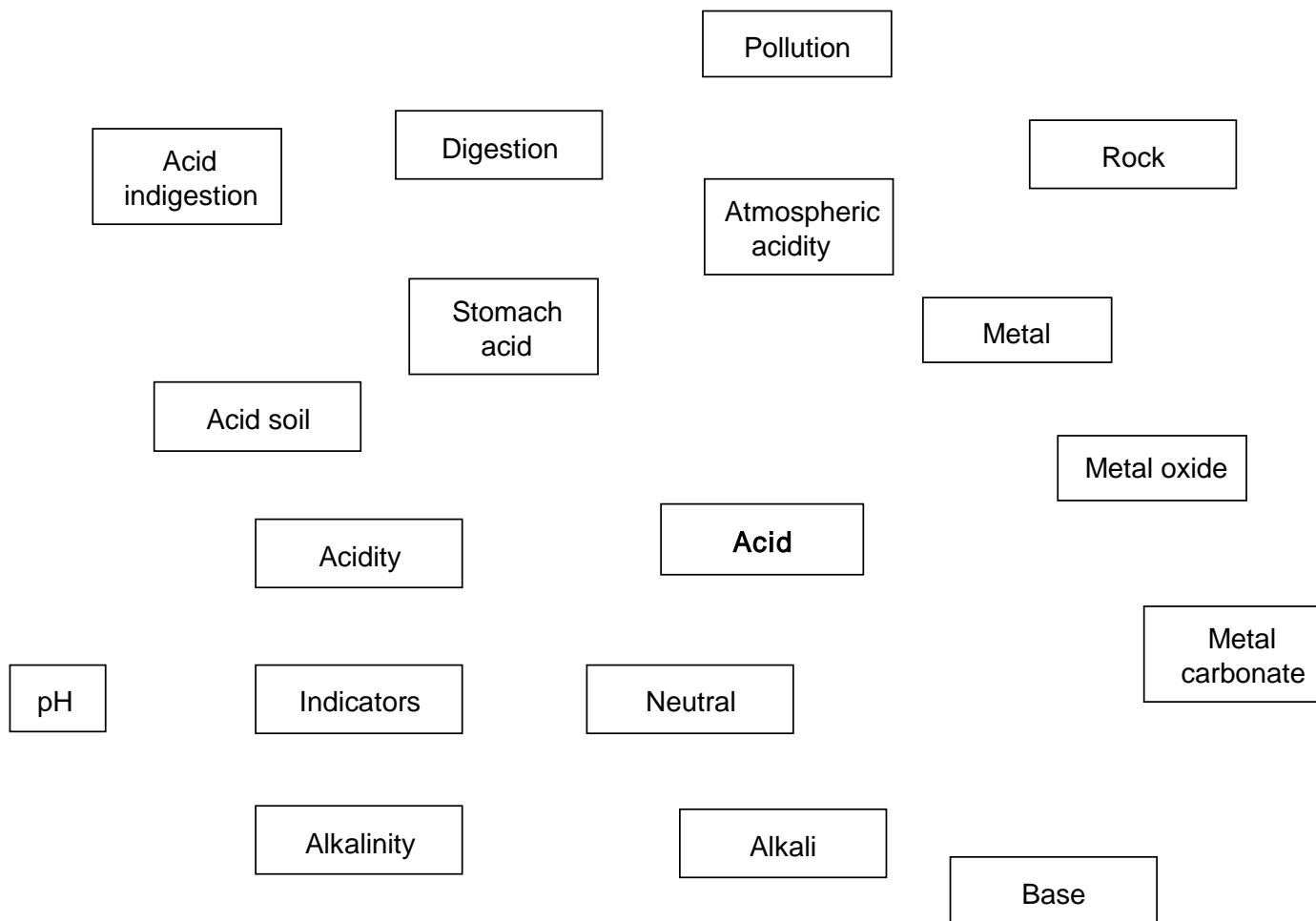
The links are not explained on the map. Read through the statements below, and work out which link on the map each sentence is about. Label each line on the map with the

number of the statement – eg 7

1. Acidity is a property of acids.
2. Acids can be identified using indicators.
3. Acidity can be measured using the pH scale.
4. Acidity can be detected using an indicator.
5. Alkalinity is a property of alkalis.
6. Alkalinity can be detected using an indicator.
7. Alkalinity can be measured using the pH scale.
8. Neutral solutions can be identified using indicators.
9. Alkalis can be identified using indicators.
10. Acids are not neutral solutions.
11. Alkalis are not neutral solutions.
12. pH may be found using universal indicator.
13. Acids react with alkalis to give a salt and water.
14. Bases react with acids.
15. An alkali is a base which dissolves in water.
16. Metal carbonates are bases.
17. Metal carbonates react with acids to give a salt and carbon dioxide.
18. Metal oxides are bases.
19. Metal oxides react with acids to give salts and water.
20. Some metals react with acid to give a salt and hydrogen.
21. Acids in the air cause atmospheric acidity.
22. Atmospheric acidity is increased by some forms of pollution.
23. Atmospheric acidity causes weathering of rocks.
24. Pollution can increase the rate of weathering of rock.
25. Atmospheric acidity causes the corrosion of some metals.
26. Pollution can increase the rate of corrosion of metals.
27. Acid is found in the stomach.
28. Stomach acid helps us digest our food.
29. Too much stomach acid can cause indigestion.
30. Some bases are used to relieve acid indigestion.

- K. Some _____ contain too much acid for many plants to grow.
- L. _____ is sometimes added to soil to neutralise acidity.
- M. Acids react with _____ to give a salt and water.
- N. Bases react with _____.
- O. An _____ is a base which dissolves in water.
- P. Metal carbonates are _____.
- Q. _____ react with acids to give a salt and carbon dioxide.
- R. Metal oxides are _____.
- S. Metal oxides react with _____ to give salts and water.
- T. Some _____ react with acid to give a salt and hydrogen.
- U. Acidity is a property of _____.
- V. Acids can be identified using _____.
- W. Acidity can be measured using the _____ scale.
- X. Acidity can be detected using an _____.
- Y. Alkalinity is a property of _____.
- Z. Alkalinity can be detected using an _____.
- α Alkalinity can be measured using the _____ scale.
- Ω Neutral solutions can be identified using _____.
- δ Alkalis can be identified using _____.
- Φ Acids are not _____ solutions.
- Σ Alkalis are not _____ solutions.
- ψ pH may be found using universal _____.

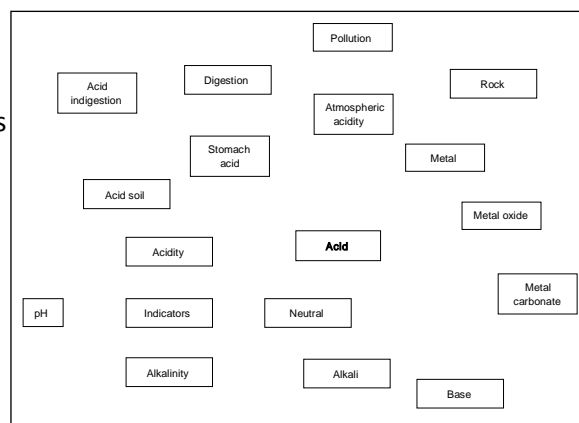
Outline acid revision map



Connecting up the revision map

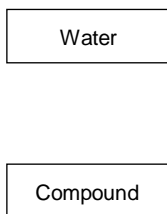
You have been given a copy of an outline of a revision map for the topic of acids. This shows some of the things you may have met when you studied acids and bases in your science class. However the map is not complete!

The boxes on the map need to be connected to show how the ideas are linked.

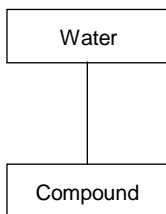


Instructions

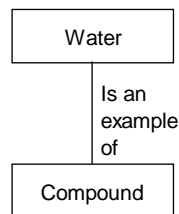
1. Look at the outline map. Find two boxes that you think you can connect.
2. Draw a clear line between the two boxes.
3. Add a label to the line to explain the connection.



Step 1



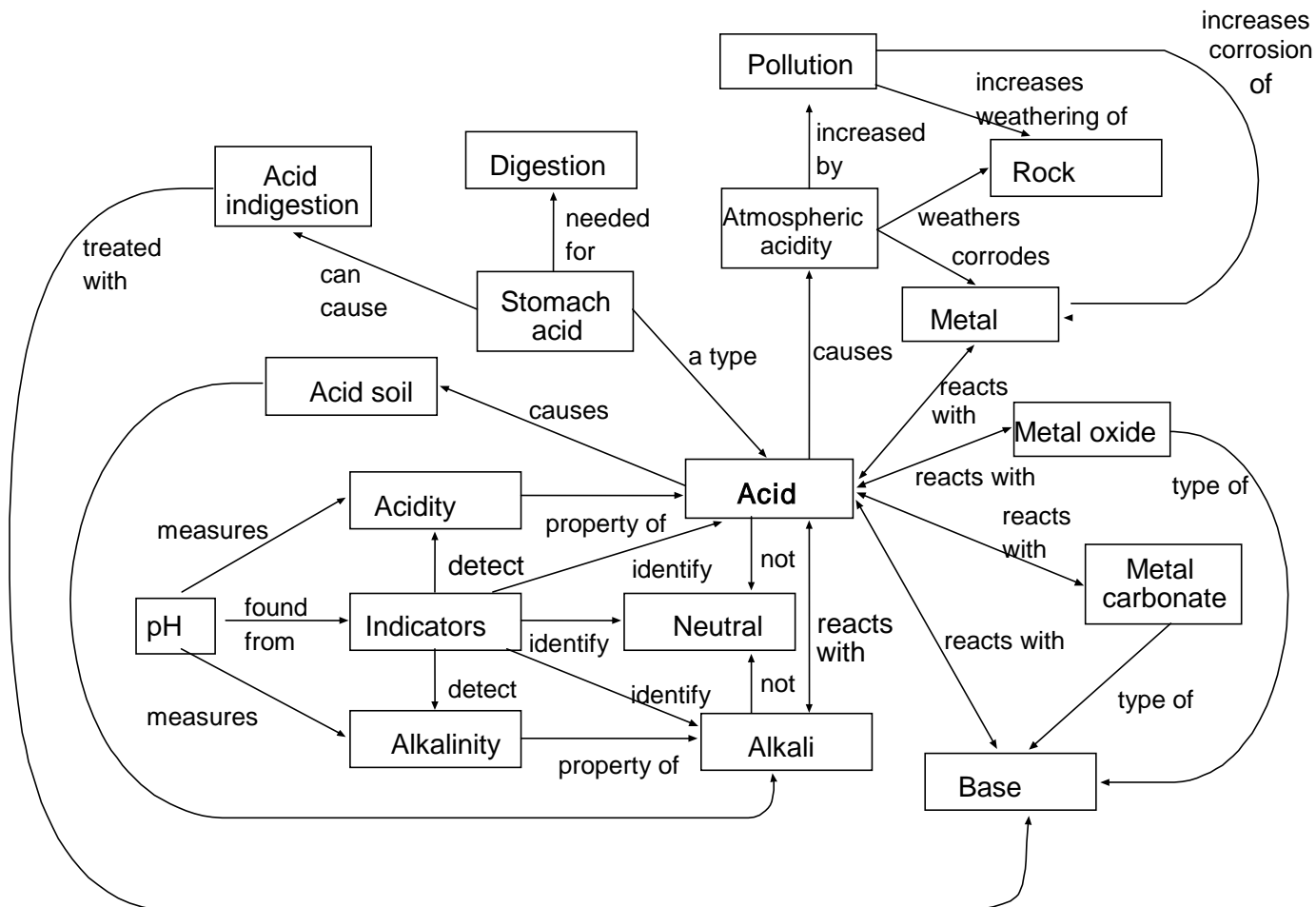
Step 2



Step 3

4. Repeat for as many connections as you can find.
5. See if you can think of any other boxes that would fit on this revision map. Draw them in.
6. Show the connections for the new boxes in the same ways as above (steps 2 and 3).

Example concept map – acids



Useful Websites

A collection of activities about acids and alkalis, including red cabbage indicators, secret messages, and taste tests.

<http://www.miamisci.org/ph/>

Glossary

Term	Description
Concentration	The amount of a substance within a known volume of a mixture
Dissociation	The splitting of a compound into smaller particles
Effervescent	The release of gas from a solution
Indicator	A coloured substance that changes colour with hydrogen ion concentration
Ion	An atom or molecule where the number of electrons is not equal to the number of protons
Neutralisation	A chemical reaction where an acid and base react together to form a salt
pH	A measure of the acidity or basicity of an aqueous solution
Proton donor	An acid donates protons in a reaction
Mole	A measure of the amount of a substance related to the number of atoms in 12 g of carbon-12.
Titration	A laboratory method of quantitative chemical analysis that is used to determine the unknown concentration of an identified substance

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